Name ____ AP Chemistry

EQUILIBRIUM PROBLEMS

Manipulation of Equilibrium Constant

The equilibrium constant changes based on the temperature of the reaction. However, if the reaction equation is changed at all at this temperature, then the equilibrium constant is also changed in a similar fashion. The different changes fall into two basic categories:

- Reversing the reaction

- Multiplying the reaction by a coefficient

Often reactions occur completely in the gaseous state, but sometimes scientists may wish to analyze them in terms of concentrations. Therefore, there needs to be a relationship between K_p (equilibrium constant involving solely partial pressures) and K_c (equilibrium constant involving only concentrations):

Example #1:

The following equilibrium concentrations were observed for the Haber process at 127°C:

$$3\ H_2\ (g) + N_2\ (g) \leftrightarrow 2\ NH_3\ (g)$$

$[NH_3] = 3.1 \times 10^{-2}$ M, $[N_2] = 8.5 \times 10^{-1}$ M, and $[H_2] = 3.1 \times 10^{-3}$ M

(a) Calculate the value of K at 127°C for this reaction.
(b) Calculate the value of the equilibrium constant at 127°C for the reaction:

$$2\ NH_3\ (g) \leftrightarrow N_2\ (g) + 3\ H_2\ (g)$$

(c) Calculate the value of the equilibrium constant at 127°C for the reaction given by the equation

$$\tfrac{1}{2}\ N_2\ (g) + 3/2\ H_2\ (g) \leftrightarrow NH_3\ (g)$$

Example #2:

Using $K_p = 1.9 \times 10^3$, calculate the value of K at 25°C for the reaction

$$2\ NO\ (g) + Cl_2\ (g) \leftrightarrow 2\ NOCl\ (g)$$

Calculating Equilibrium Constant and Determining Missing Equilibrium Concentrations or Pressures

Given all of the equilibrium concentrations and/or pressures, the equilibrium constant can be calculated. However not always are all of the equilibrium concentrations or pressure available, but the equilibrium constant is. Therefore, the missing equilibrium value can be determined.

Example #3:

Dinitrogen tetroxide in its liquid state was used as one of the fuels on the lunar lander for the NASA Apollo missions. In the gas phase, it decomposes to gaseous nitrogen dioxide:

$$N_2O_4\ (g) \leftrightarrow 2\ NO_2\ (g)$$

Consider an experiment in which gaseous N_2O_4 was placed in a flask and allowed to reach equilibrium at a temperature where $K_p = 0.133$. At equilibrium, the pressure of N_2O_4 was found to be 2.71 atm. Calculate the equilibrium pressure of NO_2.

Questions (Part I):

1) At 627°C, K = 0.76 for the reaction

$$2\ SO_2\ (g) + O_2\ (g) \leftrightarrow 2\ SO_3\ (g)$$

Calculate K at 627°C for (a) the synthesis of one mole of sulfur trioxide gas and (b) the decomposition of two moles of SO_3.

2) At 25°C, $K = 2.2 \times 10^{-3}$ for the reaction

$$ICl\ (g) \leftrightarrow \tfrac{1}{2}\ I_2\ (g) + \tfrac{1}{2}\ Cl_2\ (g)$$

Calculate K at 25°C for (a) the decomposition of ICl into one mole of iodine and chlorine and (b) the formation of two moles of ICl.

3) Given the following data at a certain temperature,

$$2\ N_2\ (g) + O_2\ (g) \leftrightarrow 2\ N_2O\ (g) \qquad K = 1.2 \times 10^{-35}$$

$$N_2O_4\ (g) \leftrightarrow 2\ NO_2\ (g) \qquad K = 4.6 \times 10^{-3}$$

$$\tfrac{1}{2}\ N_2\ (g) + O_2\ (g) \leftrightarrow NO_2\ (g) \qquad K = 4.1 \times 10^{-9}$$

Calculate K for the reaction between one mole of dinitrogen oxide gas and oxygen gas to give dinitrogen tetroxide gas.

4) Consider the reaction

$$2\ NO\ (g) + Cl_2\ (g) \leftrightarrow 2\ NOCl\ (g)$$

At a certain temperature, K = 1.8. Calculate P (NO) if P (Cl_2) = 0.0393 atm and P (NOCl) = 0.217 atm at equilibrium.

5) At a certain temperature, K is 0.040 for the decomposition of two moles of bromine chloride gas (BrCl) to its elements. An equilibrium mixture at this temperature contains bromine and chlorine gases at equal partial pressures of 0.0493 atm. What is the equilibrium partial pressure of bromine chloride?

6) For the reaction

$$N_2\ (g) + 2\ H_2O\ (g) \leftrightarrow 2\ NO\ (g) + 2\ H_2\ (g)$$

K is 1.54×10^{-3}. When equilibrium is established, the partial pressure of nitrogen is 0.168 atm and that of NO is 0.225 atm. The total pressure of the system at equilibrium is 1.87 atm. What are the equilibrium partial pressures of hydrogen and steam?

Determining Equilibrium Concentrations and/or Pressures Given Initial Conditions

Sometimes the equilibrium concentrations and/or pressures are not known or given, only the initial values. The stoichiometric relationships of the reaction can be used to express the concentration and/or pressure at equilibrium in terms of the initial values.

<u>Example #4</u>:

Carbon monoxide reacts with steam to produce carbon dioxide and hydrogen. At 700 K the equilibrium constant is 5.10. Calculate the equilibrium concentrations of all species if 1.00 mol of each component is mixed in a 1.00 L flask.

<u>Questions (Part II):</u>

1) At 800 K, hydrogen iodine gas can be formed from hydrogen and iodine gases. K is 59.2. What are the partial pressures of all the gases at equilibrium when the initial partial pressures of hydrogen and iodine gas are 0.450 atm?

2) The reaction of

$$CO\ (g) + H_2O\ (g) \leftrightarrow H_2\ (g) + CO_2\ (g)$$

has an equilibrium constant of 1.30 at 65°C. Carbon monoxide and steam both have initial partial pressures of 0.485 atm, while hydrogen and carbon dioxide start with partial pressures of 0.159 atm. Calculate the partial pressure of each gas at equilibrium.

3) For the reaction

$$PCl_3(g) + Cl_2(g) \leftrightarrow PCl_5(g)$$

The value of $K_c = 96.2$ at 400 K. If the initial concentrations are 0.22 mol/L of PCl_3 and 0.42 mol/L of Cl_2, what are the equilibrium concentrations of all species?

4) The reaction of iron and water vapor results in an equilibrium reaction,

$$3\quad Fe(s) + 4\ H_2O(g) \leftrightarrow Fe_3O_4(s) + 4\ H_2(g)$$

and an equilibrium constant of 4.6 at 850°C. What is the concentration of water present at equilibrium if the reaction is initiated with 6.5 g of H_2 and excess iron oxide, Fe_3O_4, in a 15.0 L container?

5) Solid ammonium iodide decomposes to ammonia and hydrogen iodide gases at sufficiently high temperatures. The equilibrium constant for the decomposition at 673 K is 0.215. Fifteen grams of ammonium iodide is sealed in a 5.0 L flask and heated to 673 K. What is the total pressure in the flask at equilibrium?

Approximation Methods with Small Equilibrium Constants

Example #6:

Gaseous NOCl decomposes to form the gases NO and Cl_2. At 35°C, the equilibrium constant is 1.6×10^{-5}. In an experiment in which 1.0 mol NOCl is placed in a 2.0 L flask, what are the equilibrium concentrations?

Questions (Part III):

1) At a particular temperature, $K = 4.0 \times 10^{-7}$ for the reaction:

$$N_2O_4\ (g) \leftrightarrow 2\ NO_2\ (g)$$

In an experiment, 1.0 mol of N_2O_4 is placed in a 10.0 L vessel. Calculate the concentrations of N_2O_4 and NO_2 when this reaction reaches equilibrium.

2) At 25°C, K = 1.6 x 10^{-5} for the decomposition of NOCl into NO and Cl_2 gases. Calculate the concentrations of all species at equilibrium for each of the following original mixtures.
   (a) 2.0 mol of pure NOCl in a 2.0 L flask
   (b) 2.0 mol of NOCl and 1.0 mol of Cl_2 in a 1.0 L flask

3) Nitrogen gas reacts with oxygen gas at high temperature to make nitrogen oxides. This is a major source of air pollution from auto exhaust. At 1500 K, the equilibrium constant for the formation of nitrogen monoxide from its elements is 1.0 x 10^{-5}. Calculate the equilibrium concentration of NO when the initial concentration of N_2 is 0.60 M and O_2 is 0.30 M.

4) Phosgene (a component in making of compact discs) decomposes by the reaction

$$COCl_2\ (g) \leftrightarrow CO\ (g) + Cl_2\ (g)$$

for which K_p = 6.8 x 10^{-9} at 100°C. If pure phosgene at an initial partial pressure of 1.0 atm decomposes, calculate the equilibrium pressures of all species.